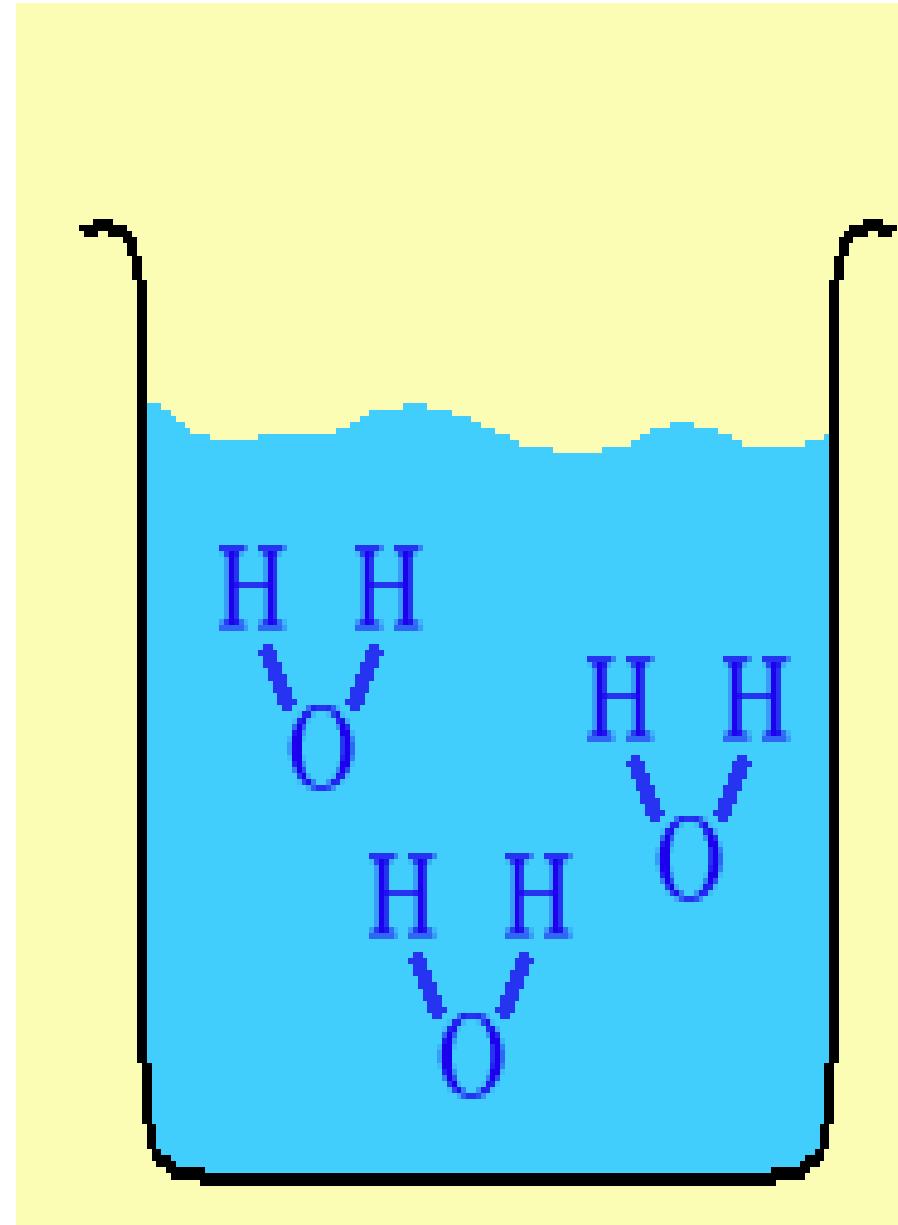


Water and Ph

Presented by :

Dr.Alaa ElTahir

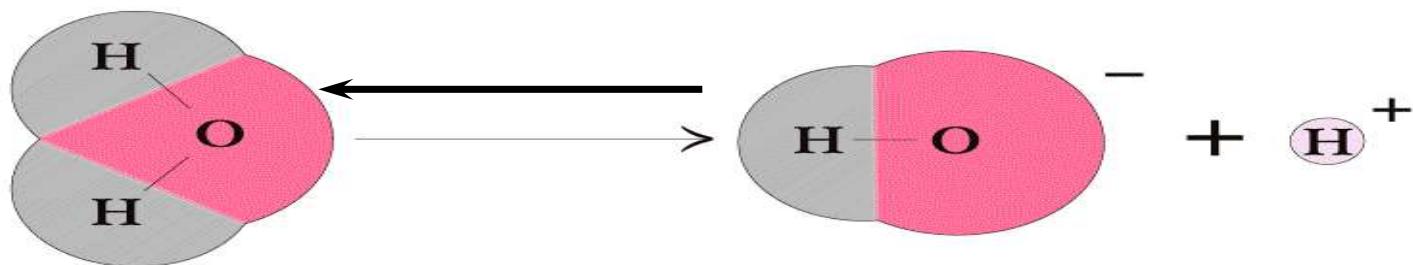


Aims:

- Chemical and Physical
- Characteristics of water
- PH
- Acid
- Base
- Acidosis
- Alkalosis
- Physiological PH
- Buffers

The Ionization of Water

Water is slightly ionized into a hydrogen ion (H^+) and a hydroxide ion (OH^-).



At room temperature, only about one out of every 10^9 water molecules is ionized at any given time.

As a result, pure water consists mostly of H_2O molecules.

Concentrations of H⁺ and OH⁻

At 25°C, water dissociates to the extent that [H⁺] and [OH⁻] are 1x10⁻⁷M. The concentration of water is a constant (55.5 M).



$$\underline{[\text{H}^+] [\text{OH}^-]} = K_{\text{eq}}$$

$$(\text{H}_2\text{O})$$

$$[\text{H}^+] [\text{OH}^-] = K_{\text{eq}} (\text{H}_2\text{O})$$

$$[\text{H}^+] [\text{OH}^-] = K_w$$

$$10^{-7} \times 10^{-7} = K_w$$

$$10^{-14} = K_w$$

The Equilibrium constant of Water

$$K_c = \frac{[H_3O^+][OH^-]}{[H_2O]^2}$$

$$K_c[H_2O]^2 = [H_3O^+][OH^-]$$

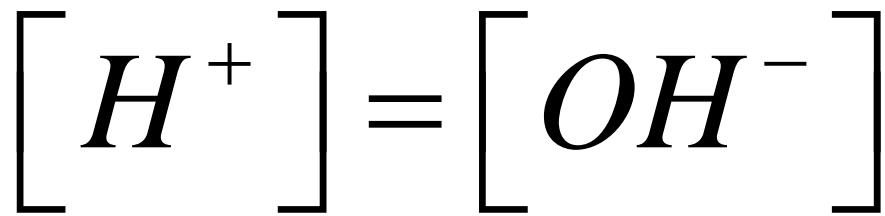
$$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14} \text{ (at } 25^\circ\text{C)}$$

$$K_w = [H^+][OH^-] = 1.0 \times 10^{-14} \text{ (at } 25^\circ\text{C)}$$

In Pure Water :

$$[H^+] = [OH^-] = 1.0 \times 10^{-7}$$

A solution for which



is said to be neutral

In most solutions, these ions are not equal in concentration. As the concentration of one increases the other must decrease so that their product equals 1.0×10^{-14}

Ion Concentration (mol/l)

	pH	Hydrogen Ion (H^+)	Hydroxyl Ion (OH^-)
Acid	0	1	0.0000000000001
	1	0.1	0.000000000001
	2	0.01	0.00000000001
	3	0.001	0.0000000001
	4	0.0001	0.000000001
	5	0.00001	0.00000001
	6	0.000001	0.0000001
Neutral	7	0.000001	0.000001
Alkaline	8	0.0000001	0.000001
	9	0.00000001	0.00001
	10	0.000000001	0.0001
	11	0.0000000001	0.001
	12	0.00000000001	0.01
	13	0.000000000001	0.1
	14	0.0000000000001	1

The concentration of

$$[H^+]$$

In an aqueous solution is usually quite small. We therefore express the concentration of hydrogen ions as the negative logarithm of

$$[H^+]$$

Osmolality

- **Osmosis**
 - Osmosis: is the movement of water from a weak solution to a more concentrated solution across a semipermeable membrane.

Isotonic, hypotonic, hypertonic: Clinical use of the terms

- **Isotonic:** same osmotic pressure as blood (same concentration of solutes)
- **Hypotonic:** lower osmotic pressure than blood (less concentration of solutes)
- **Hypertonic:** higher osmotic pressure than blood (higher concentration of solutes)

Cell size and solutions

- **Cells in isotonic solutions: ?**
- **Cells in hypotonic solutions: ?**
- **Cells in hypertonic solutions: ?**

pH

pH as a shorthand way to know how much H^+ is present in a solution. To avoid the use of large negative exponents, the pH scale was devised. pH is defined as the negative logarithm of the hydrogen ion concentration.

$$pH = -\log[H^+]$$

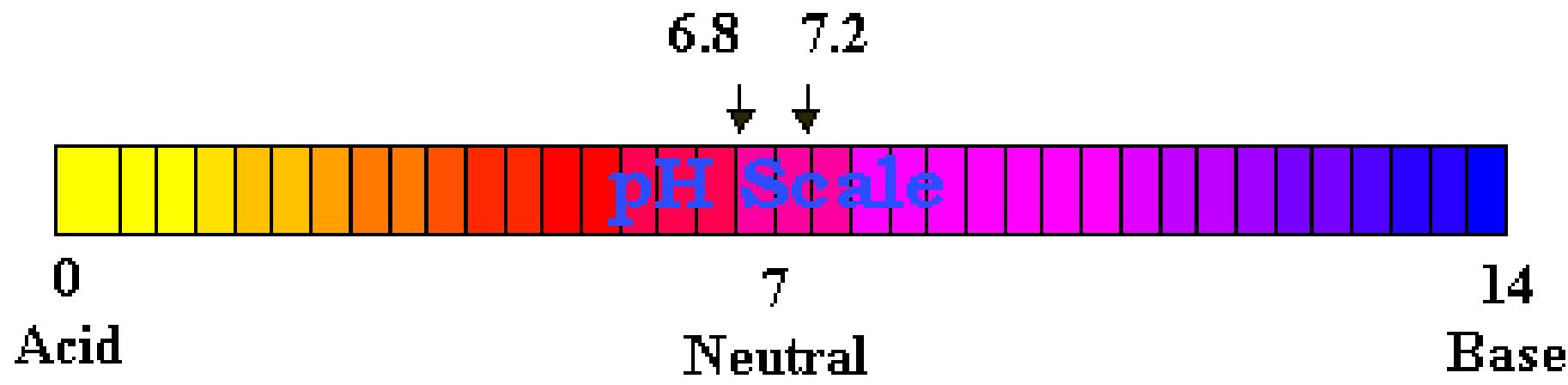
pH scale

pH scale ranges from 1 to 14

pH of 1-7 is acidic with the smaller number being more acidic

pH of 7-14 is basic with the larger number being more basic.

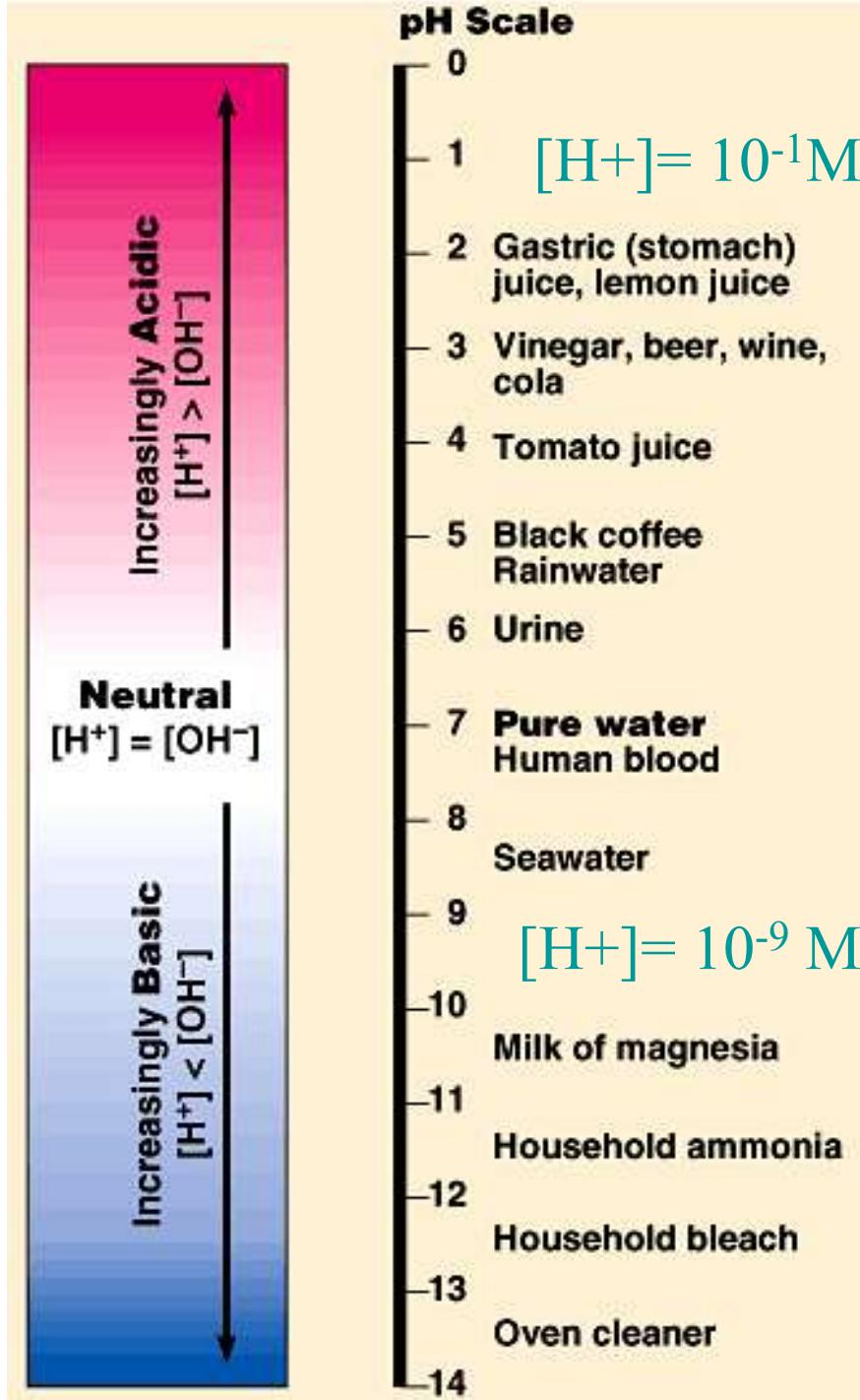
pH 7, the dissociation of water, is neutral.



pH

$$\text{pH} = -\log [\text{H}^+]$$

Log scale means 10X change per unit!

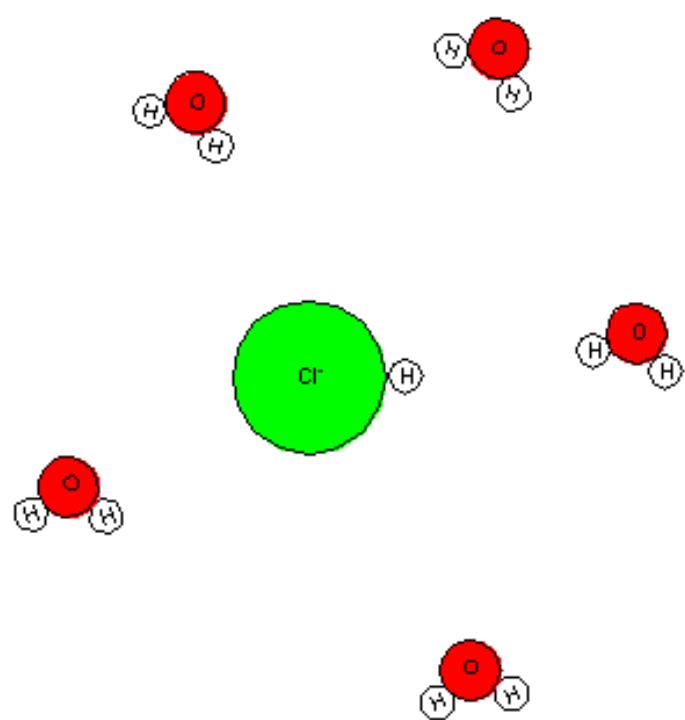


ACIDS and BASES

Definition of an acid

An acid is a proton donor.





Strong Acids

Strong acids completely ionize in water.

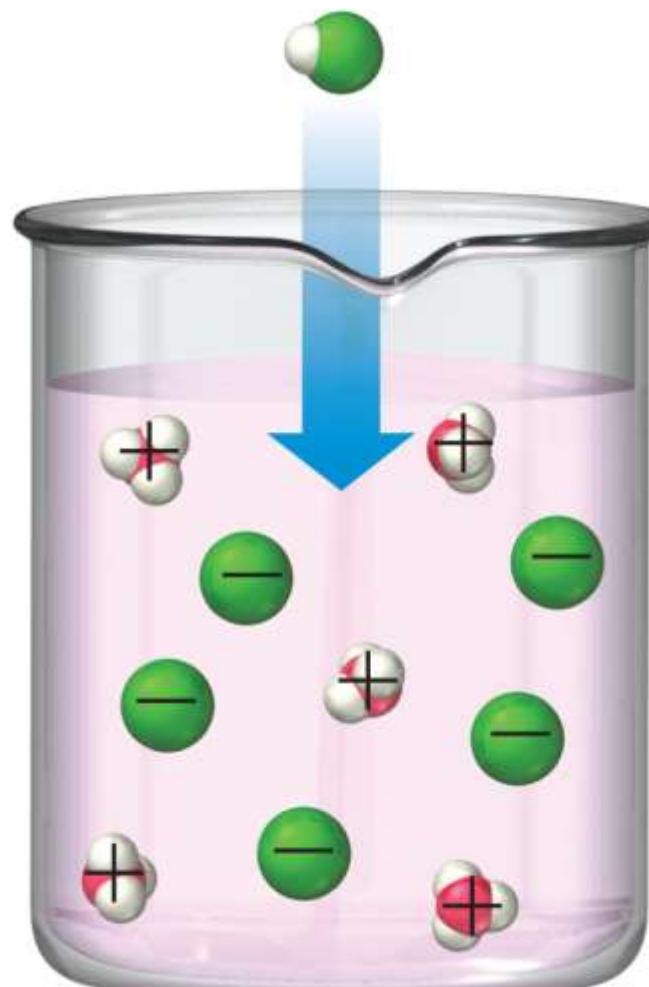
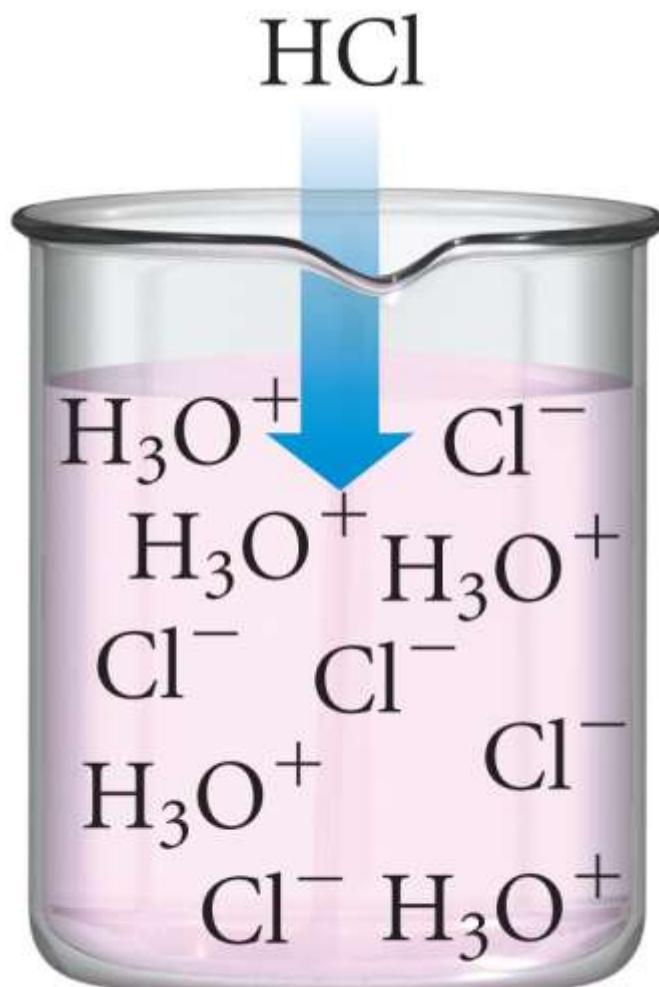
Strong acids go 100% to the right.

This is an irreversible reaction.



A Strong Acid

When HCl dissolves in water, it ionizes completely.



Weak Acids

Weak acids do not ionize completely.

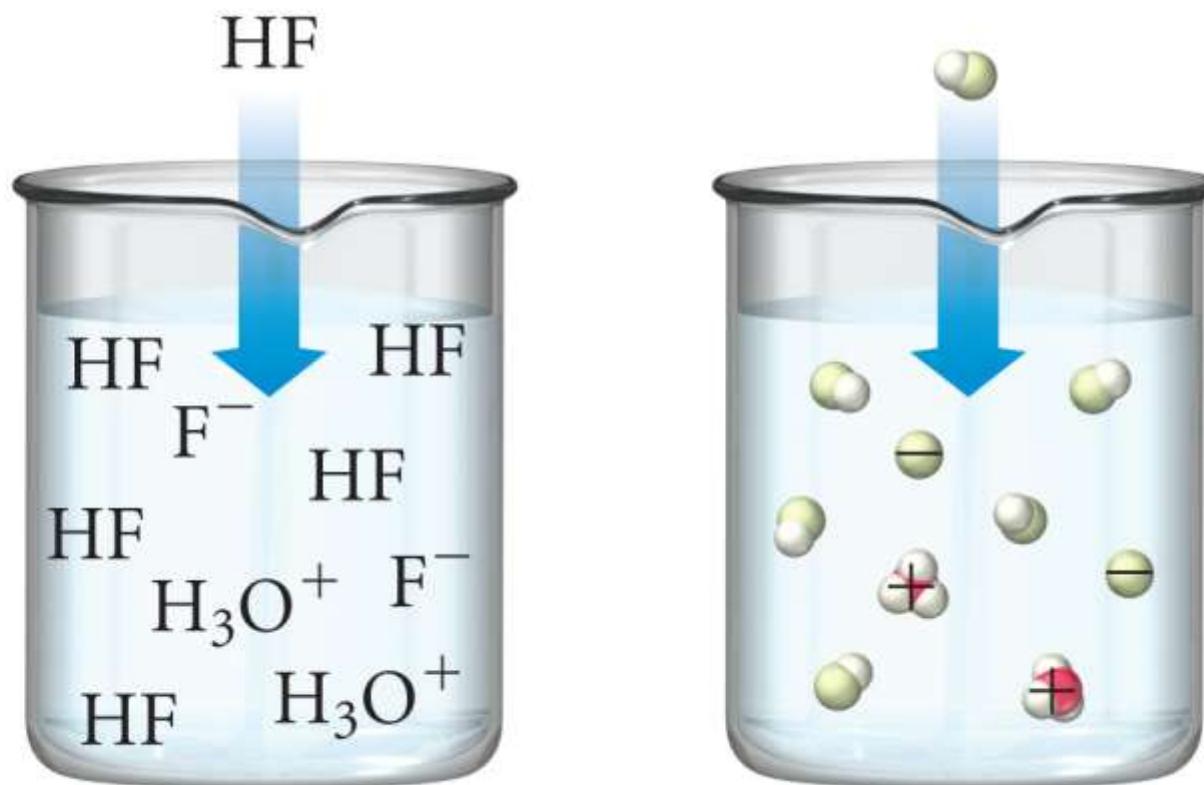
HF is a weak acid.



this is a reversible reaction.

A Weak Acid

When HF dissolves in water, only a fraction of the molecules ionize.



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Definition of a base

A base is a proton accepter or a hydroxyl donor.



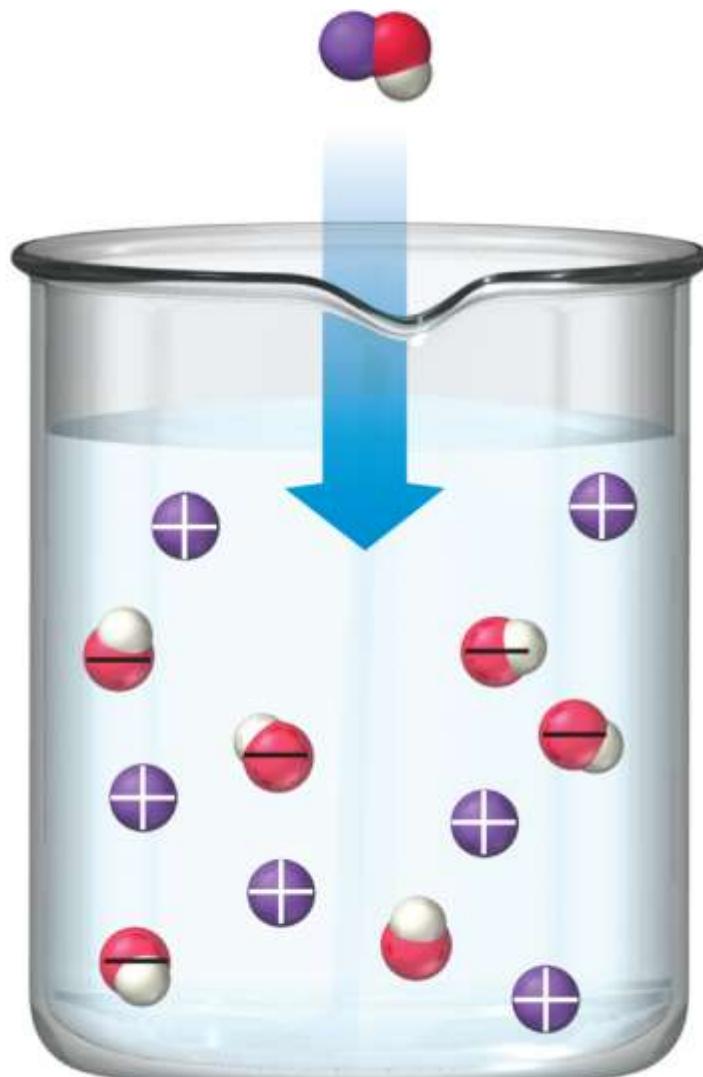
Strong Bases

A strong base completely dissociates in solution into the ions that make it up. This is an irreversible reaction.



A Strong Base

NaOH

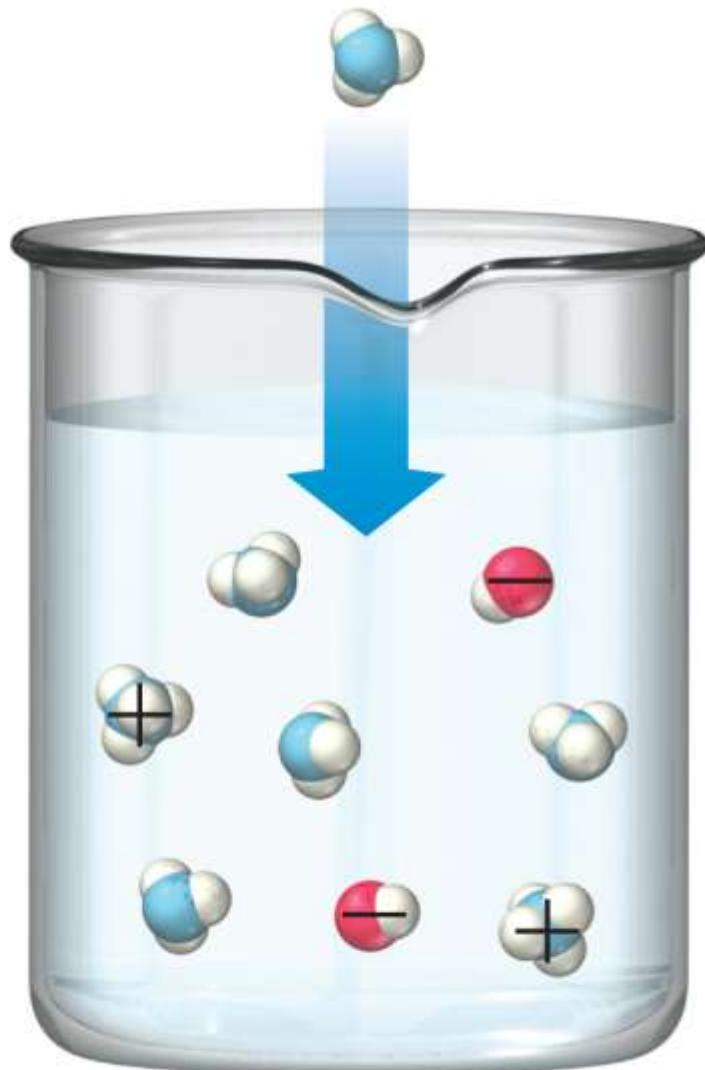
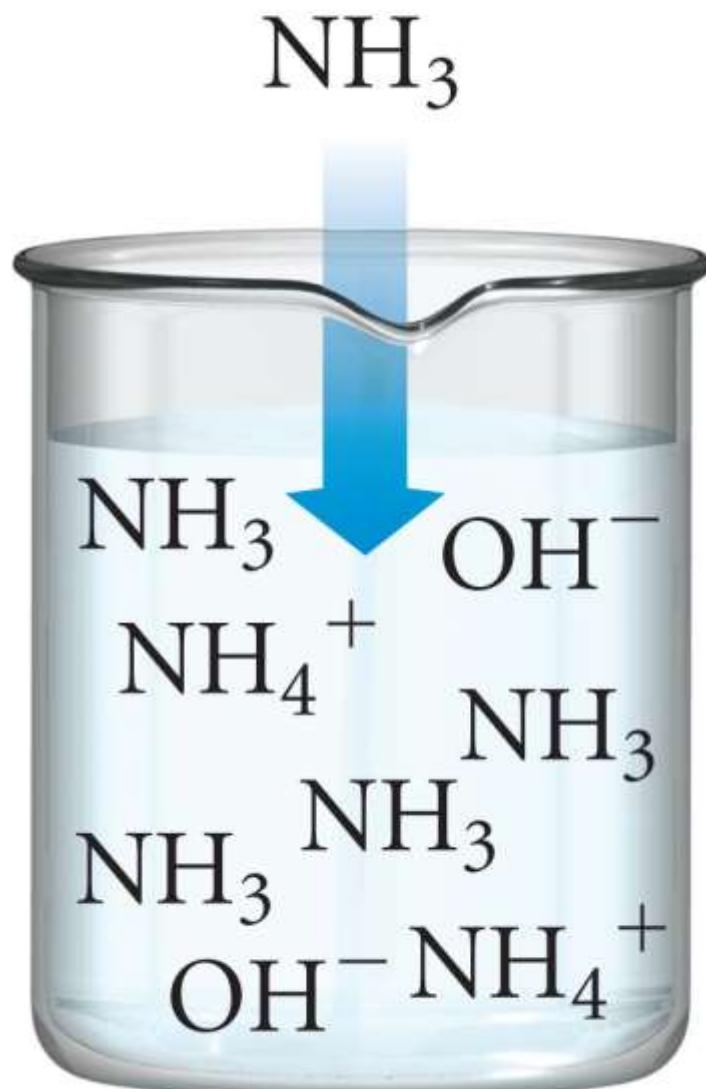


Weak Bases

Weak bases do not fully ionize in water. This is a reversible reaction.



A Weak Base



Ionization constants of Acids



$$K_a = [\text{H}^+][\text{A}^-] / [\text{HA}]$$

Ka= Measure of ability of an acid to dissociate

Strong acids: high Ka, ionized 100% into in water

Weak acids: low Ka, ionized partially (~10%) in water

The Physiological pH

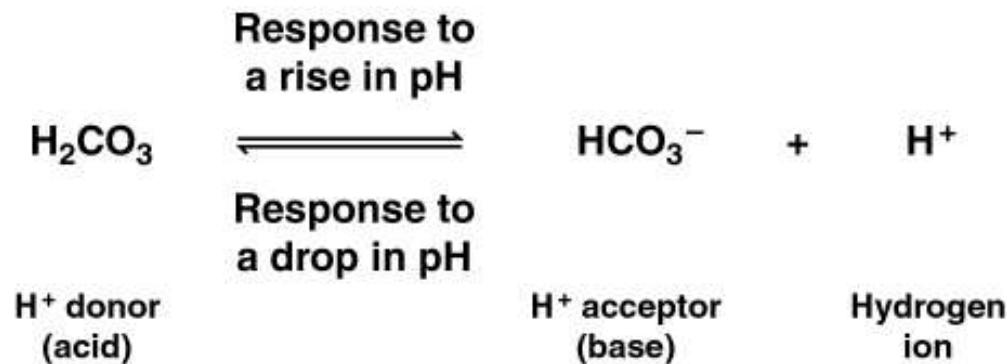
- It is important in the human body that pH remains nearly constant
 - † ~ 7 inside cells
 - † 7.4 ± 0.05 in blood
- This can be achieved by the use of **BUFFERS**.

Buffers

Are solutions of a weak acid (HA) and its conjugate base, or a weak base and its conjugate acid.

Buffers resist the change in pH because they have acid to neutralize bases and bases to neutralize acids.

- Buffers resist changes to the pH of a solution when H^+ or OH^- is added to the solution.
- Buffers accept hydrogen ions from the solution when they are in excess and donate hydrogen ions when they have been depleted.



- Two common ways to produce this effect

1. Weak acid + its *conjugate base*

e.g. $\text{CH}_3\text{COOH}/\text{CH}_3\text{COO}^-$

2. Weak base + its conjugate acid

e.g. $\text{NH}_3/\text{NH}_4^+$

- Note that buffers require an acidic component

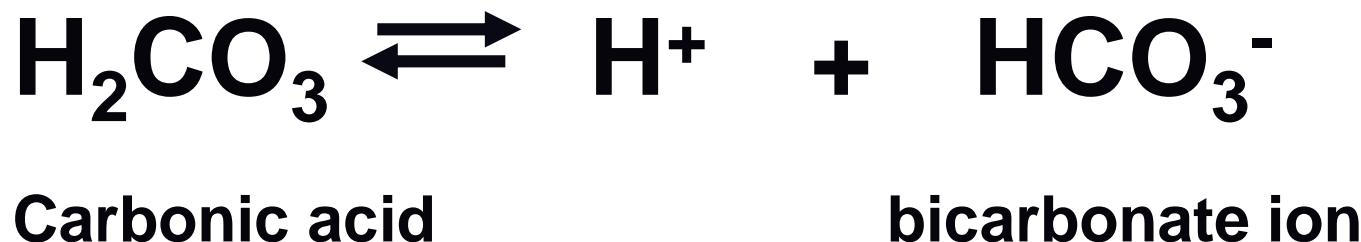
and a basic component

The buffer systems of the body

(a) Proteins are the most important buffers in the body. They are mainly intracellular and include haemoglobin.

(b) **Phosphate** buffer (H_2PO_4^- : HPO_4^{2-}) is mainly **intracellular**.

(c) H_2CO_3 is a buffer found in the blood (extracellular).



Physiological Buffers :

Acid/base balance is one of the major problems in clinical practice.

H⁺ conc. In plasma & extracellular fluid is 40 nmol/L i.e. pH 7.4 (normal range 7.35 ---7.45) pH below 7.0 & above 7.8 can have serious consequences

Regulation of blood pH is controlled by

certain physiological mechanisms

- * Blood buffers**
- * Respiratory mechanisms**
- * Renal mechanisms**

THANK
YOU